

Sports Drinks

Fighting Dehydration with Chemistry





IN THIS CHAPTER

- Solutions and electrolytes
- Concentrations of solutions
- Acidity and pH
- Solubility and precipitation
- Stoichiometry
- Limiting reactants
- Properties of solutions

INTRODUCTORY ACTIVITY

What do you think are the benefits of drinking a sports drink while exercising, rather than plain water?

How are your ideas influenced by the marketing strategies of the companies that sell these drinks?

Share your ideas with a partner and then as a class.



This chapter has the
CHEMISTRY YOU NEED TO KNOW

to create a good-tasting carbonated beverage.

Introduction

Gatorade originated on the football fields of the University of Florida—well, actually it was in the laboratories at the University. But in 1965, the football players used it to prevent dehydration in the hot, humid Florida weather. In 1967, it came on the market, named after the Florida Gators football team.

When we exercise, we burn carbohydrates. When these carbohydrates are gone, the body begins finding energy in other ways, which produces lactic acid in the muscles. Lactic acid lowers the pH of the muscles and causes muscle fatigue and cramping. It is better to have enough carbohydrates to burn than to build up lactic acid.

When carbohydrates are used in the body, the body produces heat, which can raise the core temperature of the body. Sweat cools the body by absorbing the excess heat in order to evaporate. Therefore, sweat is a necessary body function to control overheating. However, if a person were to sweat so much that he or she lost 2% of body weight, it would put stress on the heart, raise body temperature, and lower performance.



Water is just as good for you as a sports drink.

Excessive sweating also results in loss of electrolytes. Two important electrolytes in the body are the sodium cation (Na^+) and the potassium cation (K^+). These electrolytes are necessary for storing and processing carbohydrates, transmitting nerve impulses to the brain, bringing nutrients into cells, and removing waste from cells. When one or both of these electrolyte levels begin to drop due to excessive sweating, negative consequences could include muscle fatigue, weakness, heat exhaustion, and heat stroke.

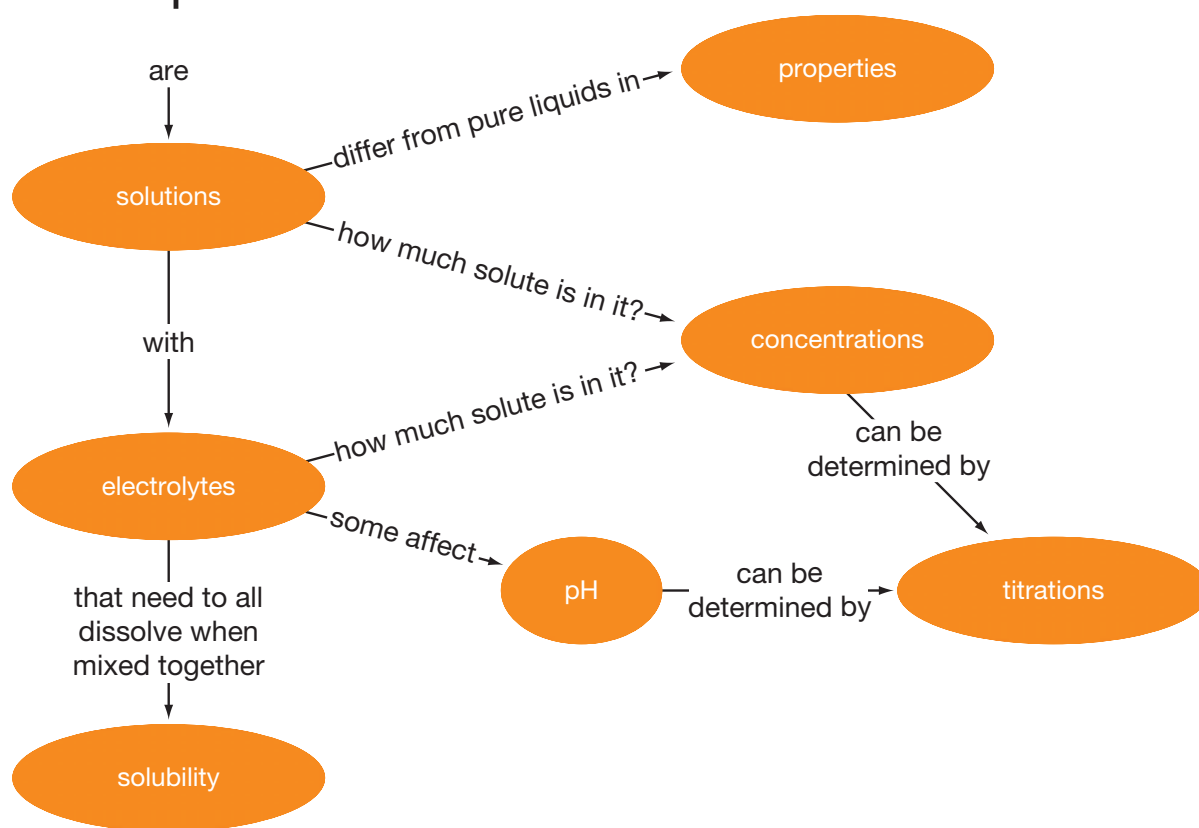
It's important to replace carbohydrates and lost water and electrolytes when exercising. But are sports drinks really better than plain water? Studies have shown that, in their ability to reduce core body temperature, heart rate, and rate of sweating, sports drinks are no better than plain water. Experts agree that it is most important to stay hydrated (and don't wait until you're thirsty to drink; you're already becoming dehydrated by that point). They also say that water rehydrates you as well as, if not better than, sports drinks. But water doesn't replace carbohydrates or electrolytes lost in sweating, you say? Those can be replaced more than adequately with a good diet before and after exercising.

So why is the sport drink industry a 700 million dollar gold mine? Because the drinks taste better! The better a drink tastes, the more likely people are to drink enough of it to stay properly hydrated. And although the sports drinks themselves don't actually improve performance, the fact that people drink more of them than they would plain water, and are therefore more hydrated, increases performance. So the sports drinks won't really increase your performance any more than maintaining a good diet and staying hydrated with water during exercise.

Electrolytes are the key to sports drinks and dehydration. But what are electrolytes, and how do they get into the drink? How do we measure how much of them are in the drink? What other reactions do they produce? And how do the properties of water change when electrolytes are added?

(Information from "Sports Drinks: Don't Sweat the Small Stuff" by Tim Graham, from the February 1999 issue of *ChemMatters*, a publication of the American Chemical Society.)

Sports Drinks



SECTION 6.1

Solutions and Electrolytes

How ionic compounds dissolve in water

In the last chapter, you learned that water is a polar molecule. Many ionic compounds dissolve in water. The polar regions of the water are attracted to the charges of the ionic compound. The positive area of a water molecule is attracted to the negative anion in the ionic compound, and the negative portion of water is attracted to the positive cation. Water molecules surround the ions on the outer layers of the solid ionic compound. Once a cation is surrounded by enough water molecules that it has a greater attraction for the water molecules than for the anions in the ionic compound, it moves off, surrounded by those water molecules. The same process happens with the anions. As these ions move off with water molecules, a new layer of ions is exposed in the solid ionic compound and the process can continue. The process of water molecules pulling away the

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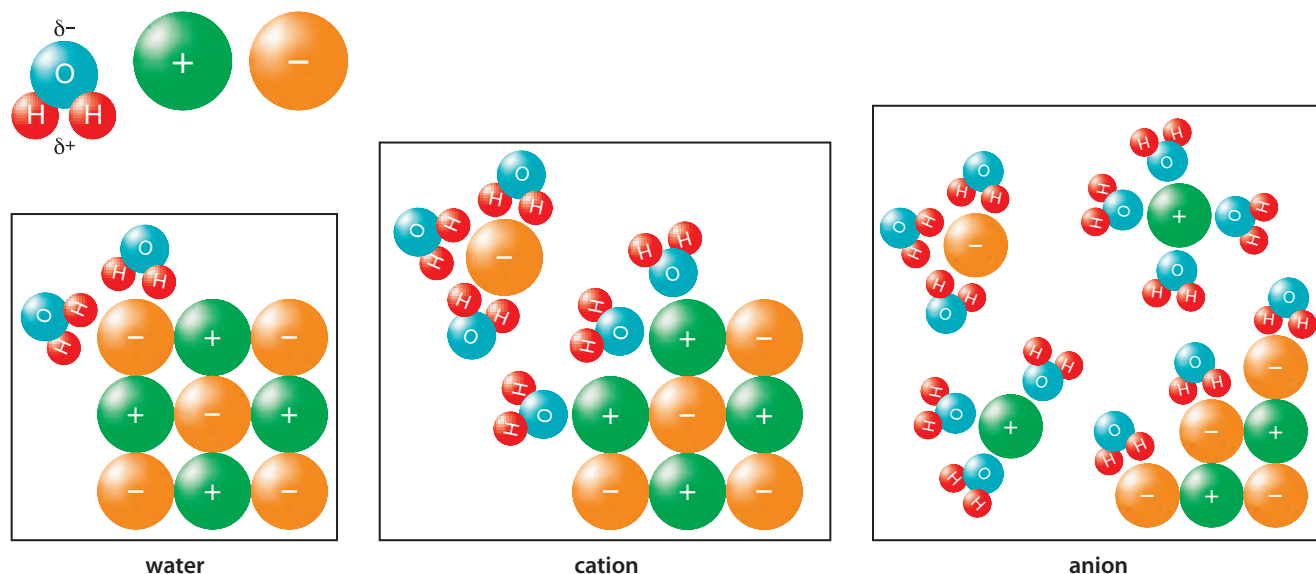


Figure 6-1. Hydration of an electrolyte. The water molecules' partially positive regions attach to anions and carry them away, while the water's partially negative regions attach to cations and carry them away. The positive and negative ions floating in water can now conduct electricity—they are **electrolytes**.

ions from an ionic compound is called **hydration**, and it is what occurs when an ionic compound dissolves in water (Figure 6-1).

A misconception about dissolving

Many people think that when a substance dissolves in water that it disappears somehow. This isn't true, of course—all the atoms are still there. However, before the substance dissolves, it is in one large mass; there are enough particles packed together that we can easily see them. As the water moves particles away from the mass one by one as described above, the particles become separate and we cannot see the individual particles anymore—they are much too tiny individually. So the substance has just been pulled into pieces that are too tiny for us to see anymore, rather than being in one large group—but they are all still there.

Forming electrolytes in water

Once the ions are hydrated, they are free-floating and move around independently of the other ions. The process of breaking up an ionic compound into the separate ions is called **ionization** or **dissociation**. Free-floating ions in water conduct electricity. When an ionic compound ionizes completely (all the ions separate), it is called a strong electrolyte, because the many free ions conduct electricity easily. An example of a strong electrolyte is table salt (NaCl). For a weak electrolyte, only a relatively few molecules ionize; most stay together. Therefore, only a few ions are available to conduct electricity, and they do so weakly. Vinegar (HCH_3COO) is a weak electrolyte.

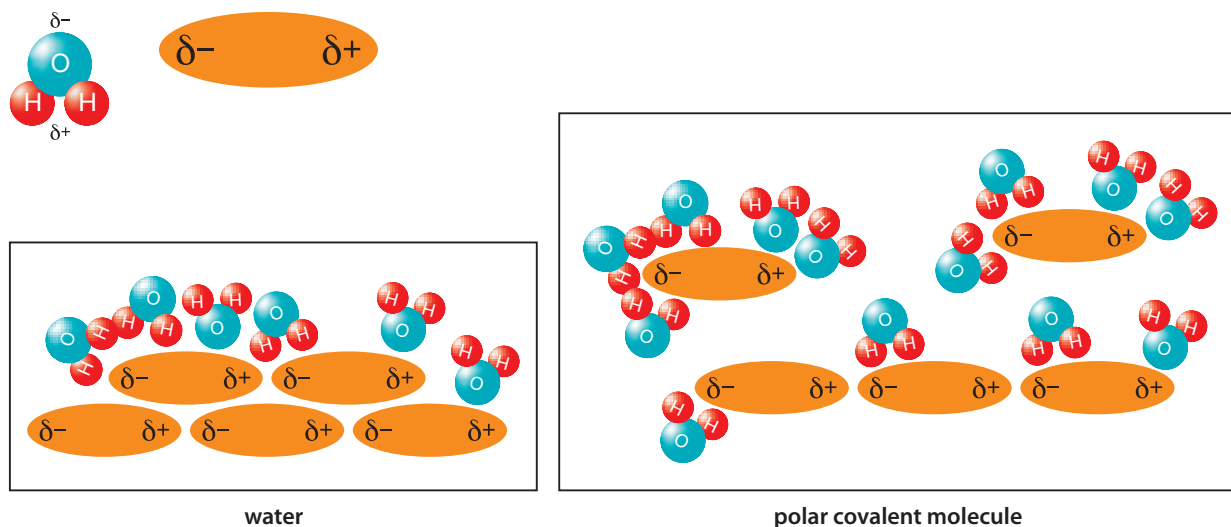


Figure 6-2. Hydration of a non-electrolyte. The water molecules' partially positive and negative regions are attracted to the regions of the opposite charge on the polar solute. Because there are no free-floating ions to conduct electricity, it is a non-electrolyte.

Non-electrolytes are substances that do not conduct electricity when dissolved in water. They do not ionize when they dissolve. Many polar covalent molecules dissolve in water, as you saw in the last chapter, but they do not conduct electricity. The hydration process occurs with these polar molecules as well—the positive portion of the water molecule is attracted to the negative portion of the polar solute molecule, and vice versa. These polar covalent molecules are separated from each other and carried away, surrounded by the water molecules. However, the polar covalent molecules are not made of ions; they share electrons with each other and therefore are not made up of cations and anions like ionic compounds. Because the polar covalent molecules do not separate into ions, dissolving them in water does not create free-floating charges that can conduct electricity. Therefore, polar covalent molecules can dissolve in water, but they do not conduct electricity (Figure 6-2). An example of a non-electrolyte is sugar ($C_6H_{12}O_6$).

Pure water also does not conduct electricity, because it is polar covalent. Tap water, however, may conduct electricity, if it has many electrolytes dissolved in it (tap water is a mixture, whereas pure water is a compound).

Clarification of terms

The terms can become confusing—“strong electrolyte,” “weak electrolyte,” or “non-electrolyte” refers to the entire compound, and describes whether it breaks up into ions when dissolved in water. However, the term *electrolyte* (as used when talking about the electrolytes in our body or sports drinks) refers to the free-floating ions, not to the whole compound from which it came.

LAB 6.1

Solution Saturation and Temperature

Purpose To observe the solubility of KBr as temperature changes

Instructions Design a lab to determine the mass of KBr that will dissolve in at least five different temperatures of water. Use no more than 10 mL of water for each trial in order to keep your usage of KBr reasonable. Your write-up must include:

- **Title**
- **Purpose**
- **Background information** (what you know about how molecules move at different temperatures, etc.)
- **Hypothesis** (make sure to include the “why”)
- **Variables and constants**
- **Materials**
- **Safety** Use caution with any hot glassware or hot plate/Bunsen burner. Wear goggles when working with chemicals and hot water.
- **Procedure**
- **Observations**
- **Data table**

- **Results/Calculations** Make a solubility curve (a graph of amount of KBr dissolved versus temperature).
- **Conclusion** Restate the purpose. Give your results. Was your hypothesis supported? If not, what new explanations can you give for your results? Explain how you think the temperature, motion of the molecules, and ability to dissolve are related, and why. What are at least two possible sources of error in this lab? How could this type of information or investigation be used in the real world?



Making a solution

Solutions are made from a solvent and a solute. The solvent does the dissolving, whereas the solute is dissolved. This section has used water as the solvent. Water is often called the “universal solvent” because it is used to dissolve so many things.

When a solution has dissolved all the solute that it can hold, it is called a **saturated** solution. The solvent cannot hold even one more molecule of solute—if any more were added, it would remain a solid. If it is holding less than the maximum that can dissolve, the solution is called an **unsaturated** solution. In an unsaturated solution, if more molecules of the solute are added, they will dissolve into the solution until the maximum of solute particles is reached and the solution is saturated.

It is possible to get a solution to hold more than the maximum amount of solute at a certain temperature. For most substances, increasing the

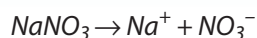
temperature will increase the amount of solute that dissolves. (There are a few substances that dissolve more in cooler temperatures, but those are far less common.) If the saturated solution is warmed, it no longer contains the maximum amount of solute—the maximum has been raised as the temperature has been raised, and it is now an unsaturated solution. More solute can be added to the solution at the higher temperature. When the solution is cooled back down again, it now has more solute particles dissolved than it ordinarily would have been able to hold at the lower temperature. This is called a **supersaturated** solution.

Example 6.1

Electrolytes

Break the following strong electrolytes into ions.

1. NaNO_3



2. CaCl_2



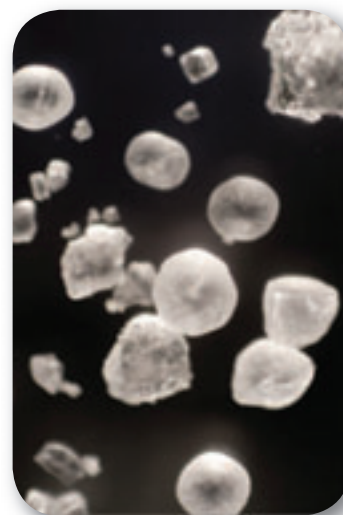
PRACTICE 6.1

Solutions and Electrolytes

1. Sodium is added to sports drinks in the form of sodium chloride (table salt). The salt is added to provide the sodium electrolyte that is lost through sweating. However, a major reason that it is added is to get someone to drink even more—salty foods make people thirsty! Break down NaCl into electrolytes.
2. Break down the following strong electrolytes: KNO_3 , Na_2SO_4 , $\text{Ca}(\text{CH}_3\text{COO})_2$, LiBr .
3. Use particle visualizations to show how the electrolytes in question 2 break down.
4. In Section 2.8, kinetics and collision theory were discussed. Increasing surface area was said to increase the rate of reaction. Explain why, at the molecular level, increasing the surface area of the solid will increase the rate of dissolving.
5. Sun tea is made from placing a see-through container of water with tea bags in a sunny window or on a porch. The sun heats the water, and the tea brews. The tea is then refrigerated and served as iced tea. Explain why adding sugar to the sun tea during or immediately after removing it from the sunny place would result in a different-tasting drink than adding sugar right before you drink it.

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Salt crystals under a microscope.

Concentrations of Solutions

There are several ways to measure the concentration of a solution. **Concentration** describes how much solute is in how much solvent—the ratio of solute to solvent (Figure 6–3).

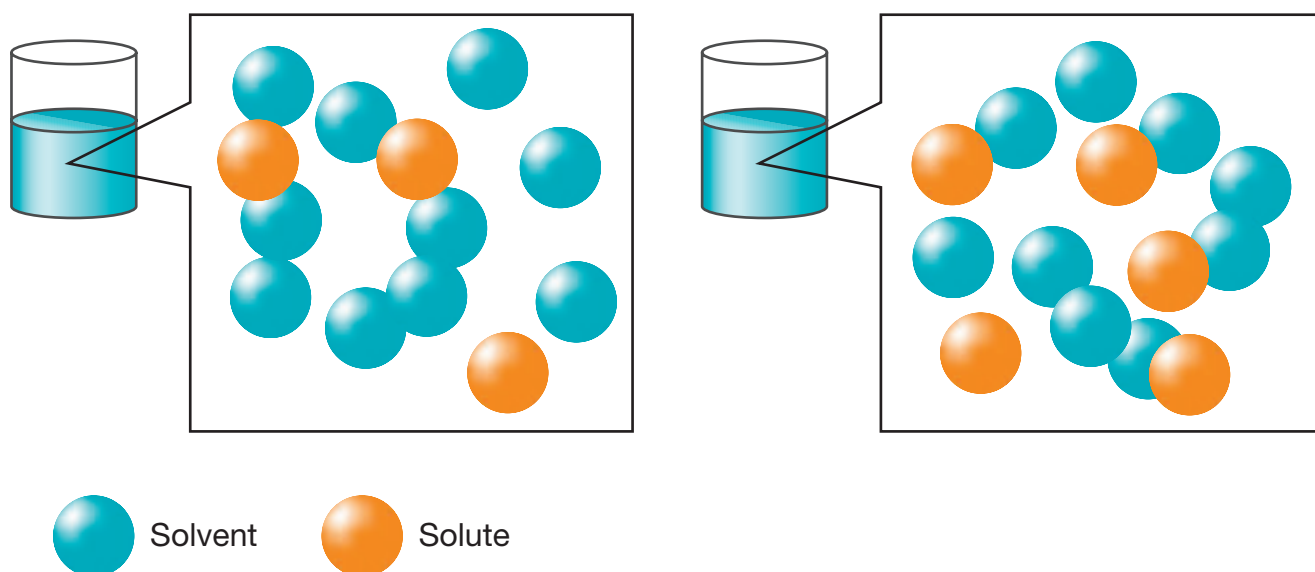


Figure 6–3. Concentration. The solution on the left has less solute in the same volume of solution. The solution on the right has more solute. The solution on the right has a higher concentration.

Percent concentration

Concentrations of solutions are often reported in percent weight per volume, %(W/V). Percent weight/volume is the percentage of the weight (more correctly, mass; although they are not scientifically the same thing, our language often interchanges the two) of the solute to the volume of the solution. The mass is in grams, and the volume in milliliters.

$$\%(W/V) = \frac{\text{mass solute}}{\text{volume solution}} \times 100$$

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Example 6.2A

Concentration in %(W/V)

Find the %(W/V) of a sample that has 14.5 g solute in 150 mL of solution.

$$\%(\text{W/V}) = \frac{\text{mass solute}}{\text{volume solution}} \times 100$$

$$\%(\text{W/V}) = \frac{14.5 \text{ g}}{150 \text{ mL}} \times 100 = 9.67\%(\text{W/V})$$

Molarity concentration

The %(W/V) unit is useful for making solutions because mass and volume are both easily measured in the lab. However, chemical equations don't use masses to balance, they use molecules or moles. Another unit for showing concentration that uses moles is molarity (M) (remember in these calculations that 0.001 L = 1 mL).

$$\text{Molarity} = \frac{\text{mole solute}}{\text{L solution}}$$

Example 6.2B

Concentration in Molarity

Find the molarity of a solution that has 26.5 g NaCl dissolved in 500 mL of solution.

Change grams to moles for NaCl:

$$26.5 \text{ g NaCl} \times \frac{1 \text{ mole NaCl}}{58.44 \text{ g NaCl}} = 0.453 \text{ mole NaCl}$$

Change mL to L:

$$500 \text{ mL} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = 0.500 \text{ L}$$

Find molarity:

$$\text{Molarity} = \frac{\text{mole solute}}{\text{L solution}}$$

$$\frac{0.453 \text{ mole NaCl}}{0.500 \text{ L}} = 0.906 \text{ M NaCl}$$

Converting between %(W/V) and molarity

%(W/V) uses mass in grams of the solute and volume in mL of the solution. Molarity uses moles of solute and liters of solution. We can convert back and forth between them because we can convert between grams and moles of solute, and we can also convert between milliliters and liters of solution.

Example 6.2C

Converting Between % and Molarity

A 750 mL sample is 0.55 M CaCl_2 . Find the %(W/V).

For %(W/V) we will need the grams solute and the mL solution.

We already have the mL solution, so we'll need to use the molarity to find the grams solute.

Use molarity and volume to find moles.

First the volume needs to be in liters to be used with molarity: $750 \text{ mL} = 0.750 \text{ L}$

$$\text{Molarity} = \frac{\text{mole solute}}{\text{L solution}}$$

$$0.55 \text{ M CaCl}_2 = \frac{x \text{ mole CaCl}_2}{0.750 \text{ L}}$$

Solve for moles:

$$(0.750 \text{ L}) \times (0.55 \text{ M CaCl}_2) = 0.41 \text{ mole CaCl}_2$$

Change grams to moles:

$$0.41 \text{ mole CaCl}_2 \times \frac{110.98 \text{ g CaCl}_2}{1 \text{ mole CaCl}_2} = 46 \text{ g CaCl}_2$$

Find %(W/V)—remember to use mL solution!

$$\%(\text{W/V}) = \frac{\text{mass solute}}{\text{mL solution}} \times 100 =$$

$$\frac{46 \text{ g CaCl}_2}{750. \text{ mL}} \times 100 = 6.1\% (\text{W/V})$$

Concentration of electrolytes

When ionic compounds dissolve and form electrolytes, the concentration for the individual electrolytes can be different from the concentration of the overall compound. You must take into consideration how the compound breaks up into electrolytes.

Example 6.2D

Concentration of Electrolytes

A solution is 0.55 M CaCl_2 (a strong electrolyte). Find the concentration of the ions.



For every one CaCl_2 , there will be 1 Ca^{+2} ion. You will get 1 Ca^{+2} mole for every 1 CaCl_2 mole. *The concentration of Ca^{+2} is 0.55 M.*

For every one CaCl_2 , there will be 2 Cl^- ions. You will get 2 Cl^- moles for every 1 CaCl_2 mole. *The concentration of Cl^- is 1.10 M.*

PRACTICE 6.2

Concentrations of Solutions

1. The minimum recommendation for carbohydrates in a sports drink is 14 grams in a 240 mL serving. What is the concentration of carbohydrates in the drink in $\%(W/V)$? The maximum recommendation is 20 grams in a 240 mL serving. What is this $\%(W/V)$?
2. Use particle visualizations to show a solution before and after more solute has been added to the container.
3. The minimum recommendation for sodium in a sports drink is 50 mg in a 240 mL serving. What is the concentration of sodium in the drink in $\%(W/V)$? (Hint: Change milligrams to grams first.) The maximum recommendation is 100 mg in a 240 mL serving. What is this concentration in $\%(W/V)$?
4. A drink that is more than 3% salt will actually dehydrate you instead of rehydrating you. The excess electrolytes are eliminated from the body by the kidneys, through producing urine. Excess urination will dehydrate the body. What is the molarity of 240 mL of a 3%(W/V) NaCl solution?
5. What is the molarity of the solution produced when 145 g of sodium chloride is dissolved in enough water to make 2.75 L of solution?
6. How many grams of potassium chloride are needed to prepare 0.750 L of a 1.50 M solution?
7. What is the molarity of the solution produced when 85.6 g of HCl is dissolved in sufficient water to prepare 0.385 L of solution?
8. To produce 3.00 L of a 1.90 M solution of sodium hydroxide, how many grams of sodium hydroxide must be dissolved?
9. If 8.77 g of potassium iodide are dissolved in sufficient water to make 4.75 L of solution,
 - a. what is the molarity of the solution,?
 - b. what is the molarity of K^+ and I^- in the solution?
 - c. what is the total ion molarity?
10. In order to prepare 2.00 L of a 3.00 M solution of iron(III) chloride, how many grams of iron(III) chloride must be used?
11. If 35.0 g $CaCl_2$ is added to 250.0 mL of water, what is the concentration in
 - a. $\%(W/V)$?
 - b. Molarity?
 - c. What are the molarities of Ca^{+2} and Cl^- in the solution?
 - d. What is the total ion molarity?

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LAB 6.2

Concentration of Solutions

Problem To find the concentration of a CuCl_2 solution; the concentration is to be found in molarity and in % (W/V)

Instructions You will be given 50 mL of a CuCl_2 solution with unknown concentration. Design a lab to find the concentration of the solution that includes the following elements. (Remember that it sometimes helps to think about these problems backward—what quantities are needed to solve for the two concentration units? How can you design a procedure and data table to measure those needed quantities?)

- **Title**
- **Problem**
- **Background information** (about concentrations and the two units of concentration you'll use)

- **Materials**



Safety Wear goggles. Use caution with glassware, especially hot glassware and Bunsen burners or hot plates. Use caution with solution. Report any spills or skin contact to your teacher, and immediately begin washing areas of skin contact with soap and water.

- **Procedure**
- **Data table**
- **Results/Calculations**



- **Conclusion** Restate your problem, give your results, give two possible sources of error (be specific) and one way that this type of technique could be used in real life.
- **Create a class histogram of results.** (See Lab 2.7, Discussion question 5 for instructions.) Describe the class results. Is there a clear mode? If not, give possible reasons.

SECTION 6.3

Acidity and pH

In Chapter 2, pH was measured for various household substances. pH was also used in the final lab project to judge the effectiveness of antacids. But what does pH really measure?

Calculating pH from concentration of hydronium

You learned in Chapter 2 that acids donate an H^+ ion in water to form H_3O^+ (the hydronium ion); see Figure 6–4. A strong acid is one in which

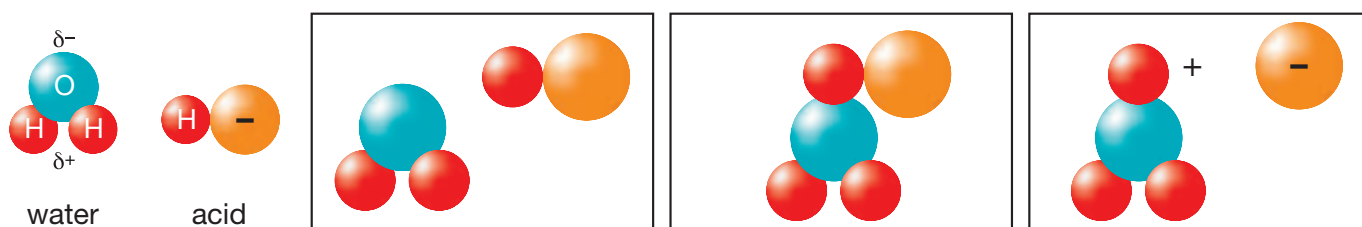


Figure 6-4. Acid dissociating in water. From left to right, the water and acid approach each other; the partially negative portion of the water molecule is attracted to the partially positive portion of the acid (the hydrogen atom). The acid dissociates, leaving an H_3O^+ hydronium ion and the anion from the acid.

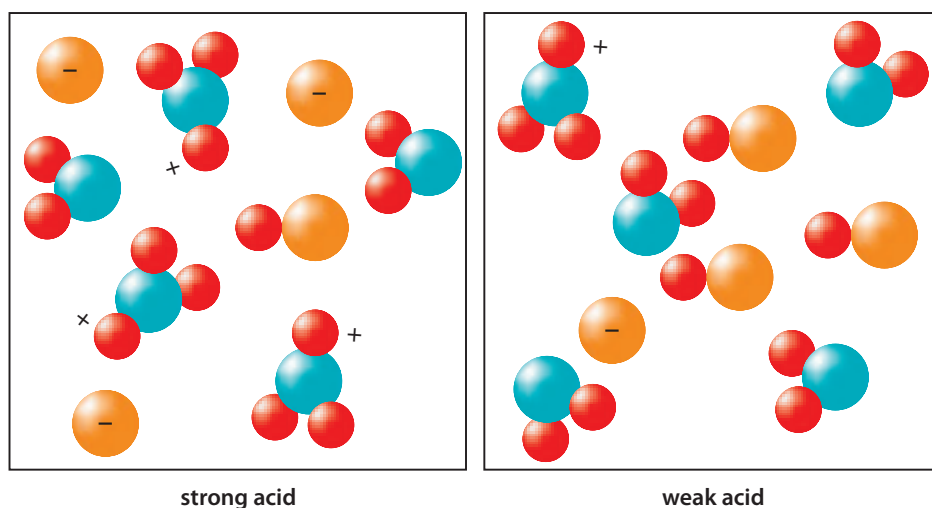


Figure 6-5. Strong versus weak acid. On the left is a strong acid; most of the acid molecules have dissociated. On the right is a weak acid; only a small portion of the acid molecules have dissociated in the water.

every molecule of the acid donates its hydrogen cation. A weak acid has only a few acid molecules that donate the hydrogen cation; the rest do not dissociate (Figure 6-5).

The “p” is a mathematical symbol that stands for “ $-\log$.” A **logarithm** (**log**) is an exponent that helps us to compute. For example, a log helps you to measure the pH or acidity of a chemical solution. The pH is the negative logarithm of the concentration of free hydrogen ions, so the pH is the “ $-\log \text{H}^+$.”

The pH is the $-\log$ of the concentration of the hydronium ion in water. The concentration is in molarity, and is symbolized by $[\text{H}_3\text{O}^+]$ in the pH equation. As $[\text{H}_3\text{O}^+]$ increases, the pH decreases (Figure 6-6).

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

The log scale is based on 10's; $\log_{10}(x) = y$, where y is the power to which 10 is raised to obtain x . In other words, $10^y = x$. For example, $\log_{10}(100) = 2$ because $10^2 = 100$.

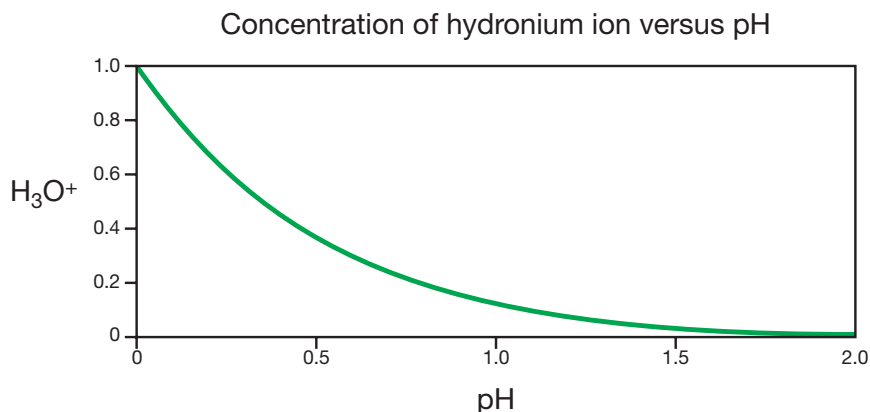


Figure 6-6. When the concentration of hydronium ion is high, the pH is low, and vice versa.

The difference between each pH value means there is a difference in the concentration of the hydronium ion of 10 times ($10\times$). For example, there are $10\times$ as many hydronium ions in a solution that has a pH of 3 as in a solution with a pH of 4. Figure 6-7 shows the relationship between acidity and pH.

Log scales have no units—pH has no units.

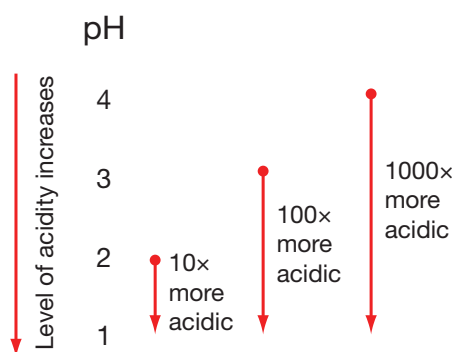


Figure 6-7. This chart shows how the level of acidity increases with each decrease in pH.

Solving for concentration of hydronium

Often the pH of a sample is known, but the concentration of the hydronium ion is what is needed. The pH equation can be rearranged algebraically to solve for the concentration of the hydronium ion.

$$[H_3O^+] = 10^{-pH}$$

Strength of acids and pH

Because most of the molecules of a strong acid produce hydronium ions in water, you don't need as many of the strong acid molecules to begin with before the concentration of hydronium starts to climb and the pH goes down. However, in a weak acid, relatively few molecules dissociate and produce hydronium ions—a lot of weak acid molecules must be added to the solution before very many of them produce hydronium. It takes a high concentration of a weak acid to lower the pH by the same amount as a small amount of strong acid.



The Beckman pH-meter, Model G, was truly a revolution in pH measurement. Originally introduced as a way to measure the acidity of lemons, chemists soon realized its potential and began using it around the country. What made the Beckman G so revolutionary in the 1930s was its compact, self-contained system of electronics and electrodes that allowed chemists to easily and quickly determine the pH of almost any solution.

Bases

As you learned in Chapter 2, bases produce the hydroxide ion (OH^-) in water. As for the acids, a strong base is one in which all the molecules of the base produce hydroxide ions. A weak base has only a few of the base molecules producing hydroxide ions in water (Figure 6–8).

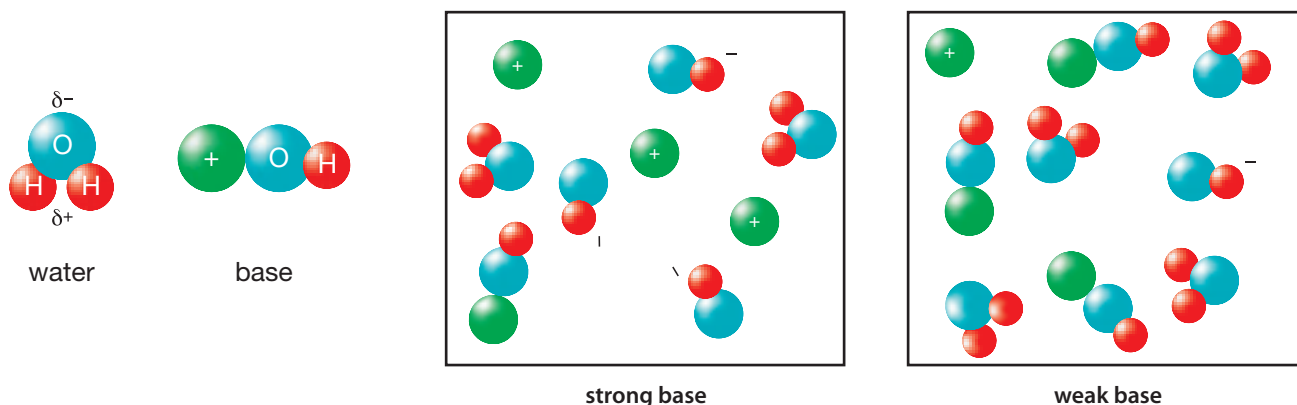


Figure 6–8. Strong versus weak base. The left shows a strong base; almost all the base molecules have dissociated in water. The right shows a weak base; a very small portion of the base molecules have dissociated.

Acids and bases are electrolytes

A strong acid is one that donates its hydronium ion to water; this means that it is dissociating the hydrogen cation from the anion. This forms electrolytes, as discussed in section 6.1. A strong acid is one that dissociates all its molecules. Therefore, a strong acid is a strong electrolyte, whereas a weak acid is a weak electrolyte.

Bases form the hydroxide ion in water; this means that bases are also electrolytes. Just as with acids, strong bases are strong electrolytes and weak bases are weak electrolytes.

Autoionization of water

Water will autoionize—it will split into ions on its own ($2 \text{ H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{OH}^-$). Water will do this until the product of the hydronium ion concentration and the hydroxide ion concentration in a solution is equal to 1.0×10^{-14} at 25°C .

$$[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14}$$

If the concentration of hydronium is greater than the concentration of hydroxide, the sample will be acidic. Likewise, if the concentration of hydroxide is greater than that of hydronium, the sample is basic. Both will be present because of the autoionization of water, but one may be in a higher concentration than the other. If the concentration of the two ions is equal, the solution is neutral, and the pH is 7.

Example 6.3

pH Calculations

1. Find the pH of a solution with a hydronium concentration of 1.5×10^{-4} .

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.5 \times 10^{-4}) = 3.82$$

2. Find the hydronium concentration of a solution with a pH of 10.7.

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-10.7} = 1.99 \times 10^{-11} \text{ M}$$

3. Find the hydroxide concentration of a solution with a pH of 12.5.

First find the hydronium concentration:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-12.5} = 3.16 \times 10^{-13} \text{ M}$$

Then use the hydronium concentration to find the hydroxide concentration:

$$1 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-],$$

$$1 \times 10^{-14} = [3.16 \times 10^{-13} \text{ M}][\text{OH}^-],$$

$$\frac{1 \times 10^{-14}}{3.16 \times 10^{-13}} = [\text{OH}^-] = 0.0316 \text{ M}$$

PRACTICE 6.3

pH

1. What would you expect the pH of a lemon-lime flavored drink to be—above or below 7? Why?
2. The pH of a sports drink can be as low as 2.38. (Such an acidic drink can be harmful to your teeth!) What is the hydronium concentration in such a sample?
3. How many times more acidic is a solution with a pH of 2 compared to one with a pH of 6?
4. Use particle visualizations to draw a solution with a pH of 2 versus a solution with a pH of 3.
5. Use particle visualizations to draw a solution with a pH of 8 versus a solution with a pH of 7.
6. If pH of the body drops below 6.8 or rises above 7.8, death can occur. What are the hydronium ion concentrations for these pH's?
7. What are the corresponding hydroxide concentrations for questions 2 and 3?
8. Find the pH if the $[\text{H}_3\text{O}^+]$ is:
(a) 1.5×10^{-2} M; (b) 0.15 M; (c) 0.000045 M; (d) 2.9×10^{-4} M.
9. Find the $[\text{H}_3\text{O}^+]$ if the pH is:
(a) 11.5; (b) 1.5; (c) 7.0; (d) 8.2.
10. Find the $[\text{OH}^{-1}]$ of the following:
(a) pH = 8.9; (b) $[\text{H}_3\text{O}^+] = 0.000035$ M; (c) pH = 2.6; (d) $[\text{H}_3\text{O}^+] = 2.65$ M.

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SECTION 6.4

LAB 6.4

Solubility and Precipitation

Purpose To observe precipitation reactions and determine solubility rules

Materials Solutions of electrolytes, well-plate



Safety You will be using very small amounts in this lab, but you still need to use caution with the solutions. Silver nitrate (the Ag^+ solution, and the

product in several of the other reactions) will stain your clothing and skin a black/brown color. Chromates (in the CrO_4^{2-} solution) are known carcinogens. Wear goggles at all times. Report spills or skin contact to your teacher, and immediately begin to wash areas of skin contact with soap and water.

(continues)

LAB 6.4

Solubility and Precipitation (*cont'd.*)

Procedure

1. It may help to place the well-plate on top of a white piece of paper so that you can see the colors of solutions and precipitates better.
2. Place 4–5 drops of each solution in the wells on the well-plate to match the data table below. The same anion is added in every well in a column, and the same cation is added in every well in a row.
3. Record observations about whether a precipitate forms, and, if so, what color it is.
3. Which anion(s) were the least likely to form a precipitate? Which were the likeliest? What about the cations?
4. AgCl is used in black-and-white photography; it is the chemical that coats the photographic paper. Use your observations in this lab to explain why you must use photographic paper in a darkroom.
5. Share your solubility rules with the class. Use a whiteboard if your teacher provides one. After sharing, do you want to make any changes to your set of rules? Were there any discrepancies? How did you use evidence to resolve them?

Data table

	Cl^-	CrO_4^{2-}	SO_4^{2-}	NO_3^-
Na^+				
Ag^+				
Ba^{+2}				
Ca^{+2}				

Discussion

1. For each precipitation reaction that occurred, write the balanced equation. Use the cation and the anion as the reactants, and the precipitate as the product. Be sure to include the charges when writing cations and anions that are not in a compound.
2. For each of the anions used in the lab, make a list of cations that it is not soluble with—the cations with which it formed a precipitate. You've created a set of solubility rules.



Solubility and Precipitation

You can't go to a shelf in the chemical storeroom and find a bottle of solid ions—the solids in those bottles are in the form of neutral compounds, not charged ions. When they are not dissolved in water, electrolytes are in the form of solid ionic compounds. The ionic compound must be neutral; therefore, if the desired electrolyte is an anion, it bonds with a cation

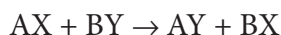
to form a neutral compound. When the desired electrolyte is a cation, it is bonded to an anion to form a neutral compound. A desired cation electrolyte and a desired anion electrolyte could bond together to form a neutral compound, as well. These compounds must be soluble in water in order to produce electrolytes.

Precipitation reactions

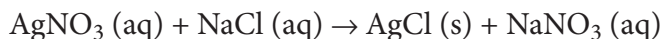
When you add these soluble ionic compounds to the water, they dissolve and produce electrolytes. However, the ions could bond with other ions in the solution and form an insoluble compound. The insoluble compound is called a **precipitate**. A precipitate is the insoluble compound that forms from the chemical reaction between two soluble salts. This reaction is then called a **precipitation reaction**. A precipitation reaction is another example of a double replacement reaction. (Acid–base neutralization was an example of a double replacement reaction discussed in Chapter 2.)

When writing precipitation reactions, it helps to include the state of matter after each chemical. This lets the reader know which compounds are dissolved in water and which compounds are precipitates. Compounds that dissolve in water are labeled “aq” for “aqueous.” The precipitates are labeled with an “s” to show that they are a solid, and are not dissolved in the water.

General form of double replacement reaction:



Example of precipitation reaction:



Solubility of ionic compounds

Why are some ionic compounds soluble in water, whereas others are not? If the attraction between the ions in the ionic compound is greater than the attraction between the water and the ions, the water cannot “pull” hard enough on the ions to cause them to separate from the other ions. In other words, the ions of the ionic compound are more stable with each other than they are when hydrated by water molecules, and therefore they don’t separate into electrolytes and dissolve.

Solubility rules

There are charts and tables that you can look at to determine whether a combination of a cation and an anion will dissolve or not. A table of solubility rules can be used when writing chemical reactions to determine whether a product will be insoluble and will form a precipitate. One such table is at the end of this chapter on page 249.

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Example 6.4

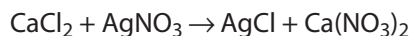
Precipitation Reactions

Finish the following equation: $\text{CaCl}_2 + \text{AgNO}_3 \rightarrow$

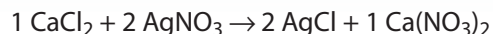
Indicate which products are aq (they dissolve in water) and which are precipitates (label as s). If nothing precipitates, write "no ppt."

First, finish the reaction by writing the products (it will be a double replacement reaction).

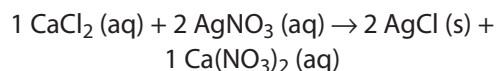
Write the formulas of the products correctly by balancing charges:



Then balance the chemical equation:



Then use the solubility rules to determine which compounds dissolve in water and which don't:



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PRACTICE 6.4

Precipitation Reactions

Finish the following reactions and balance the equations. Remember that you must write chemical formulas correctly (by balancing their charges) before you balance the equation. Indicate which products are aq (they dissolve in water) and which are precipitates (s). If nothing precipitates, write "no ppt."

1. $\text{NaNO}_3 + \text{K}_2\text{CO}_3 \rightarrow$
2. $\text{AgNO}_3 + \text{KBr} \rightarrow$
3. $\text{Cu}(\text{NO}_3)_2 + \text{NaOH} \rightarrow$
4. $\text{NaCH}_3\text{COO} + \text{NH}_4\text{NO}_3 \rightarrow$
5. $\text{K}_2\text{SO}_4 + \text{NaNO}_3 \rightarrow$
6. $\text{KI} + \text{Pb}(\text{NO}_3)_2 \rightarrow$
7. $\text{Pb}(\text{NO}_3)_2 + \text{CuCl}_2 \rightarrow$
8. $\text{K}_2\text{CO}_3 + \text{LiNO}_3 \rightarrow$
9. $\text{Ca}(\text{CH}_3\text{COO})_2 + \text{Na}_3\text{PO}_4 \rightarrow$
10. $\text{Ba}(\text{OH})_2 + \text{Na}_2\text{SO}_4 \rightarrow$

Stoichiometry

LAB 6.5A

Introduction to Stoichiometry

Purpose To use acid/base neutralization reactions to introduce the idea of stoichiometry

Background information Acids react with bases. Phenolphthalein is used as an indicator—it's colorless in acids, and pink in bases. You will put the indicator in the NaOH (base), and add acid until all of the base has reacted and the solution is neutral (colorless).

Materials 0.1 M solutions of NaOH, HNO₃, HCl, H₂SO₄, test tube, test tube rack, 3 plastic pipettes, phenolphthalein indicator

! Safety Use caution with glassware. These chemicals will cause harm. Report any spills or contact with skin to your teacher, and immediately begin washing any affected skin with soap and water.

Procedure

1. Be careful not to mix up the pipettes throughout the lab—keep them for use in only one chemical each. When adding drops, be sure that they go down to the bottom of the test tube and don't get stuck on the side.
2. Add 20 drops of NaOH solution to your test tube.
3. Add 1–2 drops of indicator to turn the basic solution pink.
4. Add HCl one drop at a time, shaking after each drop to mix the solution. Continue adding HCl slowly until the solution turns colorless. Try not to add extra acid—just enough to turn the solution colorless. Record the number of drops required.



(continues)

LAB 6.5A

Introduction to Stoichiometry (cont'd.)

5. Wash out the test tube and repeat steps 2–4 with NaOH and HNO₃.
6. Wash out the test tube and repeat steps 2–4 with NaOH and H₂SO₄.

Data

	HCl	HNO ₃	H ₂ SO ₄
Drops added			

Discussion

1. Write the balanced equation for the neutralization (double replacement reaction) between NaOH and HCl.
2. Write the balanced equation for the neutralization (double replacement reaction) between NaOH and HNO₃.
3. Write the balanced equation for the neutralization (double replacement reaction) between NaOH and H₂SO₄.
4. Determine the class average for drops of HCl, HNO₃, and H₂SO₄ that were required to neutralize the 20 drops of NaOH.
5. Because each solution had the same concentration (0.1 M), they each have the same number of molecules per drop. So we can think of the number of drops as counting the number of molecules. Use the class averages to find the ratio of HCl drops/NaOH drops. Repeat to determine the average HNO₃ drops/NaOH drops, and the average H₂SO₄ drops/NaOH drops.
6. What do you notice about these ratios and the balanced chemical equation?
7. If you had put in 30 drops of NaOH, how many drops of HCl, HNO₃, or H₂SO₄ would have been required to turn the solution neutral?
8. If the coefficients of a balanced equation tell us how many molecules of each reactant are needed to complete one and only one reaction, then we can also use them to tell moles of each chemical needed (remember, moles are a way of counting molecules in large groups). If you have 0.15 mole NaOH, how many moles HCl, HNO₃, and H₂SO₄ are needed to neutralize the reaction?

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Mole ratio in a balanced chemical equation

When compounds react, the balanced equation gives the ratio of the compounds that are reacted and the ratio of the compounds that are produced.

For example, in the equation $2 \text{AgNO}_3 + 1 \text{BaCl}_2 \rightarrow 2 \text{AgCl} + 1 \text{Ba}(\text{NO}_3)_2$, for every two molecules of silver nitrate that react, one molecule of barium chloride is reacted. And for every set of two silver nitrates and one barium chloride that react, two silver chlorides and one barium nitrate molecule will be produced. However, molecules are too tiny to count them one by one, so we use the counting unit of moles (as discussed in Chapter 3) to count molecules in larger groups. For example,

for every two moles of silver nitrate molecules and one mole of barium chloride molecules that react, two moles of silver chloride molecules and one mole of barium nitrate molecules are produced. See the table below for how these proportions might play out in different situations.



	What's being reacted		What will be produced	
	$2 \text{ AgNO}_3 +$	$1 \text{ BaCl}_2 \rightarrow$	$2 \text{ AgCl} +$	$1 \text{ Ba(NO}_3)_2$
Situation 1	2 moles reacted	1 mole reacted	2 moles produced	1 mole produced
Situation 2	4 moles reacted	2 moles reacted	4 moles produced	2 moles produced
Situation 3	1 mole reacted	0.5 mole reacted	1 mole produced	0.5 mole produced

The mole ratio from a balanced equation can be used to determine information about one of the chemical compounds from information about a different compound in the same reaction. This is called **stoichiometry**. The key ingredient for success in stoichiometry is having a correctly written and balanced chemical equation to begin with.

Example 6.5A

Stoichiometry with Moles

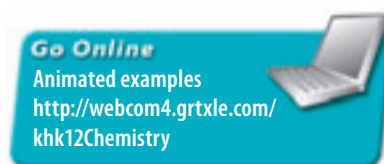
Based on the balanced equation $2 \text{ AgNO}_3 + 1 \text{ BaCl}_2 \rightarrow 2 \text{ AgCl} + 1 \text{ Ba(NO}_3)_2$, if you react 3.5 moles of AgNO_3 , how many moles of barium nitrate will be produced?

You must be in the units of mole to switch from one chemical to another. In this problem, you are in the units of mole, so you may immediately switch from AgNO_3 to $\text{Ba(NO}_3)_2$ using the ratio found in the balanced equation.

$$3.5 \text{ mole AgNO}_3 \times \frac{1 \text{ mole Ba(NO}_3)_2}{2 \text{ mole AgNO}_3} = 1.75 \text{ mole Ba(NO}_3)_2$$

Molecular mass and finding moles for stoichiometry

In Chapter 3, you learned how to convert between moles and mass using the molecular mass. Often the mass of a substance can be determined in the lab, but the moles are needed for stoichiometry. Therefore, the molecular mass is needed to convert back and forth. Remember that when you are converting between grams and moles, the molecular mass is always the number of grams equal to one mole—do not look at the balanced chemical equation when converting between grams and moles. The coefficient in the balanced equation is only used for switching between chemicals (in the unit of moles).



Example 6.5B

Stoichiometry with Masses

Based on the balanced equation $2 \text{AgNO}_3 + 1 \text{BaCl}_2 \rightarrow 2 \text{AgCl} + 1 \text{Ba}(\text{NO}_3)_2$, if you react 15.5 g of AgNO_3 , how many grams of barium nitrate will be produced?

You will need to compare one chemical to another in order to complete this problem.

You may only compare chemicals in the unit of moles. This problem has the given information in grams. Therefore, you must convert grams to moles using the molecular mass—then you can compare chemicals and complete the problem.

$$15.5 \text{ g AgNO}_3 \times \frac{1 \text{ mole AgNO}_3}{169.91 \text{ g AgNO}_3} \times \frac{1 \text{ mole Ba}(\text{NO}_3)_2}{2 \text{ mole AgNO}_3} \times \frac{261.32 \text{ g Ba}(\text{NO}_3)_2}{1 \text{ mole Ba}(\text{NO}_3)_2} = 11.9 \text{ g Ba}(\text{NO}_3)_2$$

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Molarity and finding moles for stoichiometry

Earlier in this chapter, the unit of molarity was discussed. Molarity is the ratio of moles of solute to liters of solution. Therefore, molarity can be used to convert between volume and moles. Once we know how many moles are in the solution, we can use this information for stoichiometry.

Example 6.5C

Stoichiometry with Solutions

What volume of a 0.10 M solution of NaOH is needed to completely react with 250 mL of a 0.10 M solution of H_2CO_3 ? The chemical reaction is $1 \text{H}_2\text{CO}_3 + 2 \text{NaOH} \rightarrow 1 \text{Na}_2\text{CO}_3 + 2 \text{H}_2\text{O}$.

Because the question is asking for information about the NaOH , you should begin the problem with the other chemical (H_2CO_3), which we know more about. You can use the molarity to convert

between the L and moles, but you need the volume to be in L in order to do that (so 250 mL = 0.250 L).

Once you have converted between L and moles, you can use the balanced equation to compare chemicals. Once you are at the correct chemical, you can use the molarity of NaOH to convert back to volume.

$$0.250 \text{ L H}_2\text{CO}_3 \times \frac{0.10 \text{ mole H}_2\text{CO}_3}{1 \text{ L H}_2\text{CO}_3} \times \frac{2 \text{ mole NaOH}}{1 \text{ mole H}_2\text{CO}_3} \times \frac{1 \text{ L NaOH}}{0.10 \text{ mole NaOH}} = 0.500 \text{ L NaOH}$$

Combining molecular mass and molarity in stoichiometry problems

A problem doesn't have to begin and end with the same unit. A stoichiometry problem could begin with the mass of one compound and end with the concentration of another, or vice versa.

Example 6.5D

Combination Stoichiometry

How many grams of NaOH are needed to neutralize 250 mL of a 0.10 M H_2CO_3 solution? The balanced equation is $1 \text{ H}_2\text{CO}_3 + 2 \text{ NaOH} \rightarrow 1 \text{ Na}_2\text{CO}_3 + 2 \text{ H}_2\text{O}$.

Again, the question is asking about NaOH, so begin with the information you know about the H_2CO_3 . Use the molarity to convert between the volume (changed to liters) and moles.

Then use the balanced equation to compare chemicals once you're in moles.

Finally, convert moles NaOH back to grams using the molecular mass of NaOH.

$$0.250 \text{ L H}_2\text{CO}_3 \times \frac{0.10 \text{ mole H}_2\text{CO}_3}{1 \text{ L H}_2\text{CO}_3} \times \frac{2 \text{ mole NaOH}}{1 \text{ mole H}_2\text{CO}_3} \times \frac{40.00 \text{ g NaOH}}{1 \text{ mole NaOH}} = 2.00 \text{ g NaOH}$$

Stoichiometry with gases

In Chapter 3, you learned about gas behavior and gas laws. Using the ideal gas law, you can calculate that any gas will be 22.4 L if there is one mole of particles and they are at standard temperature (0°C and 1 atm). Therefore, at STP, 1 mole = 22.4 L of a gas. This equality can also be used in stoichiometry to convert between grams and liters of gases.

If a problem asks for a volume of a gas at non-STP conditions, go ahead and use stoichiometry to find the volume at STP, and then use the combined gas law to convert the volume to the desired nonstandard conditions.

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Example 6.5E

Stoichiometry with Gases

Stoichiometry with Gases

1. How many liters of oxygen are needed to completely combust 1.5 L of methane (CH_4) at STP? The balanced equation is $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$.

Use the fact that 1 mole of gas at STP = 22.4 L to convert between L and moles.

$$1.5 \text{ L CH}_4 \times \frac{1 \text{ mole CH}_4}{22.4 \text{ L CH}_4} \times \frac{2 \text{ mole O}_2}{1 \text{ mole CH}_4} \times$$

$$\frac{22.4 \text{ L O}_2}{1 \text{ mole O}_2} = 3.0 \text{ L O}_2$$

2. How many liters of oxygen would be needed at 25°C and 1.2 atm? (See Section 3.7 for explanation of the gas laws.) Use the combined gas law.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{(1.0 \text{ atm} \times 3.0 \text{ L})}{273 \text{ K}} = \frac{(1.2 \text{ atm} \times V_2)}{298 \text{ K}}$$

The temperature must be in Kelvin to use the gas laws:

$$(25^\circ\text{C} + 273 = 298 \text{ K})$$

Solve for V_2 :

$$\frac{(1.0 \text{ atm} \times 3.0 \text{ L} \times 298 \text{ K})}{(273 \text{ K} \times 1.2 \text{ atm})} = V_2 = 2.73 \text{ L}$$

Keeping equalities straight

Use the table below to help you remember how to convert between different quantities.

To go between	Use the equality
Grams and moles	Molecular mass in grams = 1 mole
Moles and liters of a solution	Molarity in moles = 1 L
Moles and liters of a gas at STP	1 mole = 22.4 L at STP
Two different chemicals in a reaction	Coefficient ratio from balanced equation

Titration

Titration is a lab technique in which a solution of known concentration reacts with a solution of unknown concentration. The two solutions are reacted until every single molecule in the unknown solution has been exactly reacted with the known solution. There are no extra molecules of either reactant in the flask—they all have been matched up in the exact mole ratio from the balanced equation. This point is called the **stoichiometric point** or the **end point**.

An indicator is used to show when this point has been reached. If the stoichiometric point happens at a specific pH, then the indicator is a pH

indicator, as was used in Chapter 2. There are other ways of determining when a titration is over if pH isn't a factor. Sometimes the color of one of the reacting solutions itself will change once all the molecules have been reacted. Other types of indicators can tell the presence of a certain molecule. For example, iodide indicates the presence of starch by turning blue-black. Once all the iodide is reacted, a starch solution would no longer be black.

Percent yield

Because of errors in lab technique and measurements, and reactions that don't go to completion (which will be discussed in Chapter 8), reactions don't always produce 100% of what they are supposed to based on the stoichiometry. **Percent yield** = amount actually produced/amount predicted through stoichiometry if all the reactants react $\times 100$.

PRACTICE 6.5

Stoichiometry

Use the following equation for problems 1–6:



1. How many moles of nitric acid will be produced when 0.51 mole of dinitrogen pentoxide reacts?
2. How many moles of water are needed to produce 1.25 mole of nitric acid?
3. How many moles of water are needed to react completely with 0.78 mole of dinitrogen pentoxide?
4. How many grams of nitric acid will be produced when 0.25 mole of water reacts?
5. What is the percent yield of the last reaction if 24.25 g nitric acid are actually produced?
6. How many moles of dinitrogen pentoxide are needed to produce 5.25 g of nitric acid?
7. $\text{C}_6\text{H}_{14} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$.
How many grams of hexane (C_6H_{14}) would be needed to make 860.0 g of carbon dioxide?
8. What is percent yield of the last reaction if only 257 g of C_6H_{14} is produced in the lab?

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PRACTICE 6.5

Stoichiometry (cont'd.)

9. $\text{Al(OH)}_3 + \text{CaCO}_3 \rightarrow \text{Al}_2(\text{CO}_3)_3 + \text{Ca(OH)}_2$.
How many grams of calcium hydroxide would result from 250 grams of calcium carbonate?
10. $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 + \text{Pb(NO}_3)_4 \rightarrow \text{NH}_4\text{NO}_3 + \text{Pb(Cr}_2\text{O}_7)_2$.
How many grams of lead(IV) dichromate will be produced from 45.50 g of lead(IV) nitrate?
11. $\text{Li}_3\text{N} + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{LiOH}$.
What mass of lithium hydroxide will be produced with 0.38 g of lithium nitride reacting?
12. $\text{NaI} + \text{Cl}_2 \rightarrow \text{NaCl} + \text{I}_2$.
What mass of sodium chloride is produced when 0.29 g of sodium iodide is reacted with chlorine?
13. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{HOH}$.
How many mL of 0.100 M HCl are needed to react completely with 50.00 mL of 0.200 M NaOH?
14. $\text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{HOH}$.
How many mL of 0.150 M HNO_3 are needed to react completely with 150.00 mL of 0.250 M NaOH?
15. $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{NaCH}_3\text{COO} + \text{HOH}$.
How many mL of 0.200 M CH_3COOH are needed to react completely with 70.00 mL of 0.155 M NaOH?
16. $\text{HCl} + \text{Ba(OH)}_2 \rightarrow \text{BaCl}_2 + \text{HOH}$.
How many mL of 0.0500 M Ba(OH)_2 are needed to react completely with 55.00 mL of 0.450 M HCl?
17. $\text{H}_2\text{SO}_4 + \text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + \text{HOH}$.
If 10.2 mL of H_2SO_4 are needed to react completely with 75.00 mL of 0.340 M NaOH, what is the concentration of the acid?
18. $\text{H}_3\text{PO}_4 + \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + \text{HOH}$.
If 37.8 mL of H_3PO_4 are needed to react completely with 115.00 mL of 0.250 M NaOH, what is the concentration of the acid?
19. $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow \text{C}_2\text{H}_6\text{O} + \text{CO}_2$.
What mass of sugar ($\text{C}_6\text{H}_{12}\text{O}_6$) is required to produce 1.82 L of carbon dioxide gas at STP?
20. $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$.
How many liters of oxygen are necessary to combust 425 g of sulfur at STP? How many liters of oxygen would be needed if you were reacting at 25°C and 0.95 atm?
21. $\text{C}_6\text{H}_6 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}_2$.
Find the mass of benzene (C_6H_6) needed to produce 2.66 L of carbon dioxide gas at STP.

LAB 6.5B

Stoichiometry: Titration

Purpose To use solution stoichiometry to find the concentration of citric acid in a drink sample

Background information Citric acid is a tri-protic acid, meaning it has three hydrogens that are “acidic”—they react with bases. Phenolphthalein is used as an indicator—it’s colorless in acids and pink in bases. When the reaction reaches the stoichiometric point, all of the acidic hydrogens will have reacted with NaOH. The next drop of NaOH will cause the solution to be basic, and the phenolphthalein will turn pink.

Materials Drink sample, standardized NaOH, burette, flask, phenolphthalein indicator, graduated cylinder, beaker, funnel

Safety Use caution with glassware. NaOH is caustic—wear goggles at all times. Report any spills or contact with skin to your teacher, and immediately begin washing any contacted skin with soap and water.

Procedure

1. Obtain about 50 mL of the NaOH solution in a beaker.
2. Close the stopcock on the burette. With a small amount of NaOH, rinse the burette, swirling it around while tilting it to rinse the entire length of the burette. Drain the NaOH out of the stopcock to rinse it out as well.
3. Fill the burette with NaOH. Drain some out until the level of the NaOH is below the top mark on the burette. Record the exact volume of the NaOH.
4. Add approximately 30 mL of the drink solution to your flask. Before you add it to the flask, record the exact amount that is added, using the graduated cylinder.
5. Add 1–2 drops of phenolphthalein indicator—remember, it is colorless in acids and pink in bases. Swirl the flask to distribute the indicator.
6. Begin adding NaOH from the burette very slowly, swirling the flask the entire time. Continue adding the base slowly until one drop makes the solution in the flask turn pink. This indicates that you have added enough of the base to reach the stoichiometric point and then go slightly over into the basic side.
7. Continue swirling the flask for 30 seconds. If after 30 seconds of swirling, the light pink color remains, you’re done; record the volume of NaOH in the burette. (NOTE—In colorless solutions, the indicator will turn pale pink when the sample is basic; what color will it be if the pale pink was added to the original color of your drink sample? Look at the example your teacher shows, if any, to be sure that the color has changed.) If the color fades and disappears, continue adding base as in step 6 until the color remains for 30 seconds of swirling.
8. Empty and rinse the flask. If you have used only about 1/3 of the burette for the first trial, it is probably not necessary to refill it for the next



(continues)

LAB 6.5B

Stoichiometry: Titration (*cont'd.*)

trial. However, if the burette is less than about 2/3 full, refill it with NaOH until it is near the top line of the burette (you do not need to empty and rinse it, just refill). Record the exact volume of the burette.

- Repeat steps 4–7 for a total of three trials.
- Empty and rinse all glassware, and leave them as you found them.

Data

	Trial 1	Trial 2	Trial 3
Volume of drink (mL)			
Molarity of standardized NaOH			
Initial volume of burette (mL)			
Final volume of burette (mL)			

Calculations

- The reaction between citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) and sodium hydroxide produces sodium citrate

($\text{Na}_3\text{C}_6\text{H}_5\text{O}_7$) and water. Write the formula equation and balance it.

- For each of the three trials, find the volume of NaOH used in the titration. Convert these volumes to liters.
- Convert the volumes of drink sample in each trial into liters.
- Use the concentration and the volume of NaOH used in the titration to find the molarity of the citric acid in each trial.
- Find the average molarity of the citric acid in the drink sample.

Conclusion In complete sentences, restate the purpose of the lab, give your results, give two possible sources of error. Give one way in which this type of technique could be used in real life. Share your results as a class by creating a histogram (see Lab 2.7, Discussion question 5). Describe the class results. Is there a clear statistical mode? If not, give possible reasons.

LAB 6.5C

Stoichiometry: Gravimetric

Purpose To determine the concentration of sodium ions in a drink sample and compare it to the concentration advertised on the drink label

Background information Sodium ions are a desired electrolyte in sports drinks. In order to get the sodium ion into the drink, table salt (NaCl) is added. The chloride ion in table salt will precipitate with silver ions. If the amount of chloride ions that precipitates can be determined, then the amount of sodium ions can also be known.

Materials Drink sample, silver nitrate solution, graduated cylinder, 2 beakers, funnel, funnel rack (or ceramic triangle and ring), ring stand, filter paper, balance



Safety Use caution with all glassware. Wear goggles at all times. Silver nitrate causes brown spots on skin that will remain until the layers of skin are shed—avoid contact with silver nitrate.

Procedure

- Measure out approximately 100 mL of a drink sample in a graduated cylinder. Record the exact amount.
- Pour the drink sample into a beaker.
- Measure out approximately 25 mL of the 0.10 M silver nitrate solution. Add to the drink solution. The silver chloride will precipitate and form a white, cloudy-looking solid.

LAB 6.5C

Stoichiometry: Gravimetric (*cont'd.*)

4. Warm the solution for about 30 minutes—DO NOT BOIL! (You can complete steps 5–9 while waiting.)
5. Set up a funnel in a funnel rack (or in a ceramic triangle on a ring) and connect to ring stand.
6. Fold a piece of filter paper into quarters. Write your name on the filter paper in pencil.
7. Find the mass of the dry filter paper.
8. Open one side of the folded filter paper, place it in funnel, and wet it down with water.
9. Place an empty beaker under the funnel so that the tip of the funnel is inside the beaker and touching the side of the glass.
10. Decant the *cooled* solution into the funnel and allow the filtrate to drain through into the bottom beaker. Try to keep the solid precipitate in the beaker as much as you can, because it will clog the filter paper and take longer to drain.
11. When the beaker is close to empty, pour all the contents (including the precipitate) into the filter paper. Rinse the beaker out with a small amount of water, and pour the rinse water into the filter paper. Repeat the rinse a few times until all the precipitate is in the filter paper.
12. Allow the filter paper to drain.
13. Remove the filter paper, and lay on a piece of paper towel in the area designated by your teacher.
14. Clean all glassware, and leave it as you found it.

Day 2

1. Find the mass of the filter paper with the silver chloride.
2. Throw the filter paper and silver chloride in the designated trash container.

Data

Volume of drink sample _____

Concentration of silver nitrate solution _____

Mass of dry filter paper _____

Mass of filter paper and dried silver chloride _____

From the drink label, mg of sodium in one serving _____

From the drink label, mL of drink in one serving _____

Calculations

1. Write and balance the formula equation for this reaction.
2. From your data table, find the mass of silver chloride produced.
3. Use the mass of silver chloride produced to find the concentration (in molarity) of sodium chloride reacted in the drink sample using stoichiometry.
4. Write the balanced equation for the dissociation of NaCl in water.
5. When dissolved in water, sodium chloride breaks up into electrolytes (it's a strong electrolyte). Use the molarity of sodium chloride to find the molarity of sodium ions.
6. Change the molarity of sodium chloride into the %(W/V) concentration of sodium ions in the drink sample using the equation in the last question.
7. Make a list on the classroom board of the %(W/V) found for chloride ions for each type of drink sample used (if all the class used the same drink, just one list). Find the average of all the results for the type of drink that you used.
8. Figure the %(W/V) of sodium in the drink according to the information on the label.
9. Find the percent error between the average %(W/V) concentration of sodium ions and the value you figured from the bottle label.
10. Look at the ingredients on the bottle label. Is there any other possible source of Na^+ other than NaCl? How could you account for that concentration of Na^+ ?

Conclusion In complete sentences: restate purpose, give results, give two possible sources of error, and give one application of this type of technique in the real world. Share your results as a class by creating a histogram (see Lab 2.7, Discussion question 5). Describe the class results. Is there a clear mode? If not, give possible reasons.

Limiting Reactants

LAB 6.6

Limiting Reactants

Purpose To determine which reactant is limiting

Materials Solutions of hydrochloric acid and sodium hydroxide, graduated cylinder, beaker, materials to find the pH of a solution

Safety Use caution with glassware. Wear goggles at all times. Hydrochloric acid causes burns, and sodium hydroxide is caustic. Report all spills or exposure to skin to your teacher, and immediately begin washing exposed area with soap and water.

Procedure

1. Measure out approximately 20 mL of HCl in a graduated cylinder. Record the exact amount in the data table. Pour into clean beaker.
2. Measure out approximately 20 mL of NaOH in a graduated cylinder. Record the exact amount in the data table. Pour into beaker to react with HCl.
3. Find the pH of the solution using materials available.
4. Dispose of solution as indicated by your teacher, clean glassware, and leave it as you found it.

Data

Volume of acid reacted: _____

Concentration of acid: _____

Volume of base reacted: _____

Concentration of base: _____

pH of products: _____

Discussion

1. Write the balanced equation for the chemical reaction you performed.
2. If the two reactants in this reaction were present in exactly the amounts called for by the balanced equation (called **stoichiometric proportions**),

meaning that when the reaction was complete there would be none of either reactant left over, what compounds would be in the beaker at the end of the reaction? How do you know? What would be the pH of the final solution if the reactants were present in stoichiometric proportions?

3. Using your measured pH, what compounds were present in your beaker? Explain how you know.
4. Use the answer to question 3 to determine which reactant was left over in the reaction. This reactant is said to be "in excess."
5. Use the answer to question 3 to determine which reactant ran out in the reaction and limited how far the reaction could go. This reactant is called the "limiting reactant."
6. Draw a particle visualization showing what is in the container after the reaction is complete.
7. Using the volume and concentration of each of the reactants, use stoichiometry to determine the mass of salt that would have been produced if each of the reactants reacted completely (you will do two stoichiometry calculations).
8. Question 7 gave you two possibilities for how much salt would be produced. Which of those answers do you think will prove true? Explain why you chose your answer.
9. Was the limiting reactant from your calculations the same as the limiting reactant from your pH readings? Should they be the same? Why or why not?
10. Use your volume and concentration of the excess reactant and your concentration of the limiting reactant to determine what volume of the limiting reactant would have been needed for the reactants to be present in stoichiometric proportions.

Limiting and excess reactants

In titrations, the reactants are added so as to have exactly enough, and all molecules are reacted. However, not all reactions are done in this manner. Sometimes there are too few molecules of one of the reactants, and it runs out. This is called a **limiting reactant**: a reactant that runs out during a chemical reaction and limits how much product is made. The other reactant is said to be “in excess”—after the limiting reactant has run out and the reaction has stopped, there is excess of the other reactant.

This idea can be shown easily in the grocery store. When you buy a package of bratwurst, you get five in a package. When you shop for the buns, they come eight to a package. You can only make five sets of bratwurst and buns, regardless of the extra three buns. Therefore, the brats are the limiting reactant and the buns are the reactant in excess—there are some left over.



Limiting and excess reactants are concepts illustrated by grocery items. The brats (sold as five to a package) are the limiting reactant. The buns (sold as eight to a package) are the reactant in excess.

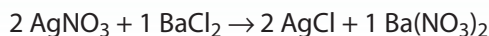


Chemistry students working on labs in 1929.

Example 6.6

Limiting Reactants

1. If 14.5 g AgNO_3 and 14.5 g BaCl_2 react, how many grams of AgCl will be produced?



Both reactants in the products have given information. Use both of these pieces of information for stoichiometry to find the mass of AgCl if each of the reactants was totally reacted.

$$14.5 \text{ g AgNO}_3 \times \frac{1 \text{ mole AgNO}_3}{169.91 \text{ g AgNO}_3} \times \frac{2 \text{ mole AgCl}}{2 \text{ mole AgNO}_3} \times$$

$$\frac{143.35 \text{ g AgCl}}{1 \text{ mole AgCl}} = 12.2 \text{ g AgCl}$$

If all the silver nitrate reacted, 12.2 g of silver chloride would be produced.

$$14.5 \text{ g BaCl}_2 \times \frac{1 \text{ mole BaCl}_2}{208.23 \text{ g BaCl}_2} \times \frac{2 \text{ mole AgCl}}{1 \text{ mole BaCl}_2} \times$$

$$\frac{143.35 \text{ g AgCl}}{1 \text{ mole AgCl}} = 20.0 \text{ g AgCl}$$

If all the barium chloride reacted, 20.0 g of silver chloride would be produced. As the silver chloride begins to be produced in the reaction, the silver nitrate and barium chloride begin to be used up. When 12.2 g of silver chloride is produced, all of the silver nitrate is gone. Once one of the reactants has run out, the reaction stops—it can't keep going without one of the reactants. Therefore, *12.2 g of silver chloride* will be produced.

2. What was the limiting reactant in question 1?

The silver nitrate is the reactant that will be totally gone once 12.2 g of silver chloride is produced—therefore, *silver nitrate is the limiting reactant*. Barium chloride will be left over once the reaction has stopped.

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PRACTICE 6.6

Limiting Reactants

1. ____ Ba + ____ $\text{N}_2 \rightarrow$ ____ Ba_3N_2 .

What mass of barium nitride is produced with 22.6 g of barium and 4.2 g of nitrogen reacted?

2. ____ $\text{Pb(NO}_3)_2$ + ____ $\text{KI} \rightarrow$ ____ PbI_2 + ____ KNO_3 .

What mass of lead(II) iodide will be produced when 16.4 g of lead(II) nitrate and 28.5 g of potassium iodide are reacted?

3. What is the limiting reactant in the previous problem?

4. ____ H_2O + ____ $\text{Na} \rightarrow$ ____ NaOH + ____ H_2 .

What is the limiting reactant when 10.0 g of water is reacted with 4.5 g of Na? (Show your work—don't just write the limiting reactant!)

5. Explain why there is no limiting reactant in a titration that is correctly done.


Properties of Solutions

LAB 6.7

Properties of Solutions

Purpose To determine how dissolved substances affect properties such as freezing and boiling point

Materials 2 beakers, test tube, 2 thermometers or temperature probes, rock salt, ice, Bunsen burner, drink sample and pure water sample, graduated cylinder

 **Safety** Use caution with glassware, especially hot glassware. Wear goggles at all times. Always point test tube away from people when heating.

Procedure

1. Thoroughly clean and rinse with distilled water two test tubes, two thermometers, and a graduated cylinder.
2. Determine the freezing point:
 - a. Label one test tube *water*. Add 10 mL of distilled water to the test tube. Place a thermometer into the test tube.
 - b. Label the other test tube *sports drink*. Add 10 mL of the drink sample to the test tube. Place a thermometer into the test tube.
 - c. Make an ice bath by adding ice, enough water to cover the ice, and rock salt to a beaker. Stir the mixture.
 - d. Place both test tubes in the ice bath—be careful not to get any of the salty ice water in them. Stir the sample tube (gently!) with the thermometer occasionally, until it freezes.
 - e. Remove the test tubes from the ice bath, and record the temperatures when reading is stable.
 - f. Do not empty the test tubes—use them for part 3.
3. Determine the boiling point:
 - a. Place the test tubes in warm water until the frozen drink sample melts. (Be careful not to get any water inside the test tubes.) Continue warming the test tubes until they no longer feel cold—do not put cold glass in hot water or a flame, or on a hot plate.
 - b. Add one or two clean boiling stones or glass beads to each test tube to promote smooth boiling.
 - c. Gently heat one of the test tubes over a Bunsen burner flame. Always point the opening in the test tube away from all people (preferably toward a wall or window). Hold the test tube with test tube tongs and move it in and out of the flame to prevent it from heating too quickly.
 - d. After the sample begins to boil, record the temperature when the reading is stable. Do not let the thermometer sit on or touch the side of the glass—take the temperature of the boiling liquid inside.
 - e. Repeat steps c and d for the other test tube.
 - f. Clean the test tubes after allowing them to cool. Be careful not to place a hot test tube into cold water.

Data

Freezing point of drink sample: _____

Boiling point of drink sample: _____

Freezing point of distilled water: _____

Boiling point of distilled water: _____

(continues)

LAB 6.7

Properties of Solutions (*cont'd.*)

Discussion

1. Classify the two samples you used as pure element, pure compound, homogeneous mixture, or heterogeneous mixture. (See Section 2.1 if you don't remember.)
2. What trend did you notice in the freezing point of a substance as solutes are added to it?
3. What trend did you notice in the boiling point of a substance as solutes are added to it?
4. Share your results from questions 2 and 3 with the class. Use whiteboards, if provided. Are there any discrepancies? If so, discuss them. How can evidence be used to resolve discrepancies?
5. When you make homemade ice cream, you put rock salt in the water surrounding the ingredients container. Use the findings from the lab to explain why you added rock salt to the ice water bath.

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Vapor pressure

The kinetic energy of molecules is related to the temperature of those molecules—the higher the temperature, the more kinetic energy they have, and the faster they move. However, temperature is related to the *average* kinetic energy of the molecules. Therefore, there are some molecules that are moving around faster and have more energy than the average, and some that are below the average.

In a liquid, there may be some molecules that are so much above the average that they have enough energy to become gas molecules. (See Figure 6–9.) If those molecules are on the top layer of the liquid—on the surface—they can “escape” the liquid, becoming a gas. That is, some of

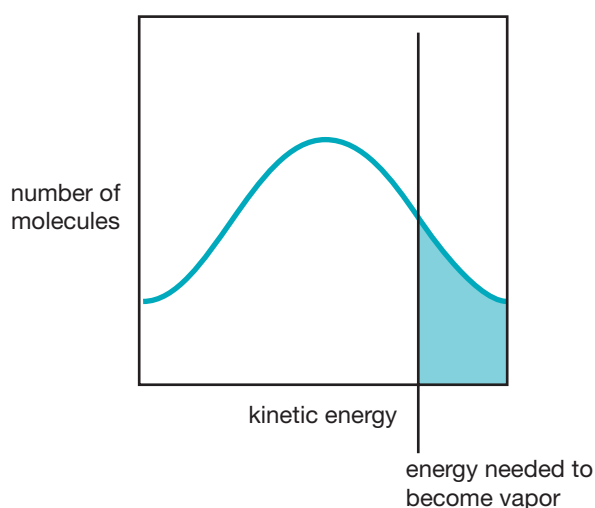


Figure 6–9. Graph showing the number of molecules that have each amount of energy. The average energy is shown, and the energy necessary to become vapor is also shown. Note that although the majority of molecules do not have the energy needed to become vapor, some do.

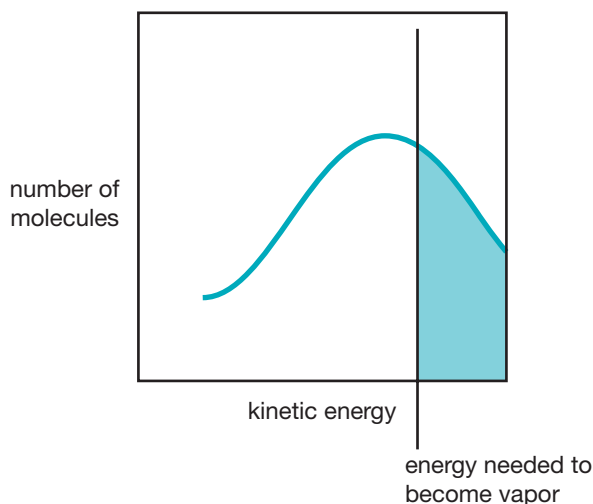


Figure 6-10. As the average energy increases, the number of molecules that have the minimum energy needed to become vapor also increases.

the molecules on the surface of a liquid have enough energy to break the intermolecular forces that hold the liquid together and pop off to become a gas. This process is **evaporation**. The gas molecules above the liquid exert pressure (you learned in Chapter 3 that all gas molecules do). The pressure of those molecules is called **vapor pressure**.

When temperature increases, the average kinetic energy increases. As the average increases, the number of molecules that have the minimum energy needed to break the forces of the liquid and become vapor also increases (Figure 6-10). Therefore, as temperature increases, vapor pressure also increases.

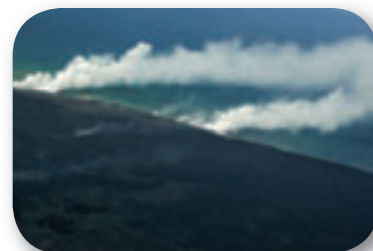
Vapor pressure of a solution

In an ionic solution, many of the water molecules are “tied up” with the solute particles. They are connected to the solute particles with intermolecular forces (see Section 5.5 for information on intermolecular forces). The water molecules evaporate less easily when they are interacting with the solute particles in this way. *The vapor pressure of a solution (with a nonvolatile solute) is always lower than the vapor pressure of the pure solvent.*

The more solute particles there are, the more water molecules are “tied up,” and therefore the lower the number of water molecules that are available to evaporate. So the more solute that is added, the lower the vapor pressure will become.

Effects of electrolyte versus non-electrolyte solutes

If the solute is a strong electrolyte, adding one molecule of the strong electrolyte compound would produce more than one solute particle. For example, adding one NaCl unit produces two particles in the water:



When hot lava from a volcano meets the cool ocean, you can see evaporation of the water as steam.



When water boils, bubbles rise to the surface.

$1 \text{ NaCl} \rightarrow 1 \text{ Na}^+ + 1 \text{ Cl}^-$. For every unit of salt that's added, you get twice the effect because you'll get two particles in the water.

Since non-electrolytes do not dissociate in solution, the effect is one to one—for every one molecule of non-electrolyte that's added, you only get one particle in the water.

Boiling point of a solution

When a liquid boils, it bubbles. Those bubbles are little pockets of molecules that have gone from the liquid phase to the gas phase. The bubbles rise because the gas is less dense than the liquid. They rise to the top until they “pop” and join with the rest of the gas above the liquid. The gas molecules in the bubbles exert the same pressure as the vapor pressure of the liquid.

If the atmospheric pressure pushing down on a sample of liquid is greater than the pressure that the bubbles would have, the bubbles cannot form. However, if the vapor pressure of the liquid increases so that the pressure of the bubbles would be at least equal to the atmospheric pressure, then the bubbles can form and the sample will boil. The way to increase the vapor pressure of the liquid is to increase the temperature. So, when the temperature increases, so that the vapor pressure increases enough that the bubbles can push against the atmospheric pressure enough to form, the solution will boil.

The vapor pressure of a solution is lower than that of the pure solvent (as discussed above). Therefore, there is a greater difference to be made up between the vapor pressure of the solution and the atmospheric pressure in order to boil the solution. This means that we have to raise the temperature even higher in order to get the vapor pressure of the solution equal to the atmospheric pressure. *Solutions always boil at higher temperatures than does a pure solvent.*

Calculating boiling point of an aqueous solution

$$\Delta T = K_B M$$

For every mole of particles that is added to a pure solvent, the boiling point increases by the **boiling point elevation constant**. For water, at sea level, the boiling point elevation constant is $0.52^\circ\text{C}/\text{mole}$. There are different constants for each solvent. This means that for every mole of particles that is added to water, the boiling point increases by 0.52° from its 100.00°C boiling point (at sea level). For 1.0 M sugar in water, the boiling point would be 100.52°C , because sugar is a non-electrolyte. For 1.0 M NaCl , which is an electrolyte producing two ions, the boiling point would be 101.04°C .

Freezing point of a solution

A substance is at its freezing point when the solid phase and liquid phase are in equilibrium. That means that every time a molecule breaks free of the solid phase and becomes a liquid phase molecule, a different liquid phase molecule joins the solid and becomes a solid phase molecule.

The overall number of solid and liquid phase molecules doesn't change. If the sample is at the freezing point, the two states of matter are in equilibrium—ice water at 0°C is in equilibrium; the amount of solid ice will stay the same. If the sample drops below the freezing point, molecules will leave the solid more slowly than they are joining the solid; eventually they will all be solid molecules—the sample completely freezes. If the sample rises above the freezing point, then molecules will leave the solid faster than the liquid molecules can join the solid; eventually they will all be liquid molecules, and the sample completely melts.

In order to understand freezing points of solution, think of an ice cube in a glass of sugar water. The solid phase ice molecules on the surface of the ice cube can break free and join the liquid phase molecules. However, the sugar molecules that are floating around in the solution get in the way of liquid phase molecules' joining the solid ice cube and becoming solid



Why does freezing preserve food and keep it from decaying? The freezing stops bacteria from growing because chemical reactions are slowed down.

phase molecules. Therefore, the molecules are leaving the solid to become liquid faster than the liquid is joining the solid—eventually all the solid molecules will leave and there will be no solid left, only liquid.

If you lower the temperature even further, the solid molecules leave the solid phase at a lower rate—a rate that can be matched by the liquid molecules joining the solid. Therefore, the rates will be at an even lower temperature for a solution than it was for a pure liquid. *The freezing point of a solution will always be lower than for the pure liquid.*

Calculating freezing point of an aqueous solution

For every mole of particles that is added to a pure solvent, the freezing point decreases by the **freezing point depression constant**. For water, at sea level, every 1 mole of particles decreases the freezing point by 1.85°C. A 1.0 M solution of sugar water (a non-electrolyte) would lower the freezing point from 0°C to −1.85°C. A 1.0 M solution of NaCl (which produces twice as many particles because it is an electrolyte) would lower the freezing point to −3.70°C.

$$\Delta T = K_f m$$

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PRACTICE 6.7

Properties of Solutions

1. Explain why vapor pressure increases as the temperature of a solution increases.
2. Explain why the vapor pressure is always lower for a solution than the pure liquid.
3. Explain why road crews put salt on the roads in winter.
4. Why would CaCl_2 have more of an effect on the roads than NaCl?
5. Explain why it is possible to get a sports drink below the freezing point of water, and still be able to drink it.
6. Why do people put salt in their water when cooking spaghetti? (Hint: It's not for taste alone! It shortens either the time to boil the water or the time to cook the pasta once the water is boiling. Which is it? Explain how you know.)
7. What concentration of particles is present in a 1.0 M solution of AlCl_3 ? What would be the boiling point and freezing point (at sea level) of a 1.0 M AlCl_3 aqueous solution?
8. What concentration of particles is present in a 2.0 M solution of CaCl_2 ? What would be the boiling point and freezing point (at sea level) of a 2.0 M solution of CaCl_2 in water?

FINAL CHAPTER PROJECT

Here is your chance to use the chemistry you know to create a good-tasting carbonated drink.

Carbonated Beverage

Purpose To use stoichiometry to make a carbonated beverage

Background information When citric acid and sodium bicarbonate (baking soda) react, they produce carbon dioxide gas, which is what “carbonates” a carbonated drink.

Materials Small paper cups for weighing ingredients, one medium paper cup for making beverage, baking soda, citric acid, powdered drink mix, measuring cup (with metric units), plastic spoon



Safety Do not do this lab in the regular laboratory, because you will be tasting the sample produced! Do not use regular lab equipment (except the balance which none of your ingredients or equipment other than the outside of a cup should touch at any time). Ingestion of excessive amounts of baking soda may be harmful.

Pre-lab calculations

Using the information on the powdered drink mix package, find the following for one serving:

1. Mass of powdered drink mix, in grams



2. Mass of sugar, in grams (1 cup of sugar = 200 grams)
3. Volume of water, in mL
4. Write the balanced formula equation for the reaction. Citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) and sodium bicarbonate react to produce carbon dioxide gas, water, and sodium citrate ($\text{Na}_3\text{C}_6\text{H}_5\text{O}_7$).
5. Use the balanced equation and stoichiometry to determine the mass of citric acid needed to react with 1.0 g of sodium bicarbonate.

Procedure

Part 1

1. Use the small paper cups and the balance to measure out the four solid ingredients needed (based on pre-lab calculations). Record observations of each of the ingredients. Add each ingredient to the large paper cup after measuring.
2. Use the measuring cup to measure the needed water.
3. Add water to the solid ingredients in the larger paper cup. Stir until all solid ingredients dissolve. Record observations during and after the reaction.
4. Taste test the drink—pour into separate cups for each lab partner. You and your partners may finish off the drink sample if you wish. Record observations about the taste.

Part 2

1. Repeat Part 1, but this time use half as much citric acid as in Part 1.
2. This time, taste test only a **small** amount of the drink.

Part 3

1. Repeat Part 1, but this time use half as much sodium bicarbonate as in Part 1.



Now that you know about stoichiometry, what do you think ecological stoichiometry is about? There is chemistry in all of life. Ecological stoichiometry studies how elements in the environment balance. The correct balance of available resources is necessary for each being's survival. The Earth's living creatures and the Earth's elemental cycles are connected. For example, how might global warming affect sea temperature, which would affect coral reefs, on which fish species depend?

2. This time, taste test only a **small** amount of the drink

Observations

Record observations of ingredients before, during, and after reaction (including taste afterward).

Discussion

1. Explain why the reaction did not occur until the water was added—why didn't they react as solids? Use your knowledge of kinetics and collision theory (from Chapter 2) and of how solutions are made (from this chapter) to answer thoroughly.
2. In Part 2, what was the limiting reactant? What was the excess reactant? How did that affect the taste?

3. In Part 3, what was the limiting reactant? What was the excess reactant? How did that affect the taste?
4. Why were you allowed to drink as much of the sample as you wanted in Part 1, but in Part 2 and Part 3 you were cautioned to taste only a small amount? (Hint: Look at the safety information.)
5. Where else would stoichiometry apply to everyday life?

(This activity adapted with permission from "Fizzy Drinks: Stoichiometry You Can Taste" by Brian Rohrig, published in *The Journal of Chemical Education*, December 2000, page 1608A.)

Strong electrolytes are ionic compounds that almost completely dissociate into ions when dissolved in water. Weak electrolytes are ionic compounds that only partially dissociate into ions when dissolved in water. When ions (which are charged) are free-floating in water, they can conduct electricity. Strong electrolytes conduct electricity well when dissolved in water, and weak electrolytes conduct electricity weakly. Non-electrolytes are covalently bonded compounds that dissolve in water but do not form free-floating charges, and therefore cannot conduct electricity.

A solution is composed of the substance that is dissolved, the solute, and the substance doing the dissolving, the solvent. A solution that still has room to hold more solute particles is called unsaturated, and a solution that is holding all the solute particles possible to hold at that temperature is saturated. A supersaturated solution is one that has been heated to a higher temperature (which generally allows a solvent to hold more solute particles) to increase the amount of solute in the solution, and then cooled down to a temperature that ordinarily would not allow the solution to hold that much solute.

The concentration of a solution describes how much solute is in how much solvent. Concentration can be given in percentage of weight/volume or molarity. The concentration of electrolytes is determined from the number of ion particles that are formed from the original ionic compound and the concentration of the original ionic compound.

pH is a measure of the acidity of a solution. $\text{pH} = -\log[\text{H}_3\text{O}^+]$. As the concentration of the hydronium ion increases, the pH decreases. Acids and bases are also electrolytes. Strong acids and

strong bases dissociate almost completely in water—therefore, they are strong electrolytes. Weak acids and weak bases dissociate only slightly, and are therefore weak electrolytes.

Not all ionic compounds are soluble in water. If two soluble compounds are mixed, it is possible that a new combination of ions that is not soluble in water will form; this is a precipitate. A series of reactions between different ions produces solubility rules that describe which ionic compounds are soluble.

Stoichiometry uses a balanced equation, with a ratio of moles of the compounds involved, to determine information about one compound involved in the reaction from information about another compound involved in the reaction. If one reactant runs out before the other reactants, it is called a limiting reactant, because the reaction stops. Stoichiometry can be used to determine the limiting reactant.

Liquids have a vapor pressure—the pressure created by those liquid molecules with enough energy to break free from the surface of the sample and become gas molecules. As temperature increases, the vapor pressure of a liquid increases. A solution always has a vapor pressure that is less than that of the original solvent, because there are fewer solvent molecules on the surface available for evaporation. A liquid boils at the temperature at which the vapor pressure of the liquid and the atmospheric pressure of the surroundings are equal. Solutions will have higher boiling points and lower freezing points than the pure solvent due to the presence of the solute particles.

CHAPTER 6

Review

- Explain the difference between a strong electrolyte, a weak electrolyte, and a non-electrolyte.
 - Define solute and solvent.
 - Explain the difference between saturated, unsaturated, and supersaturated solutions.
 - What is pH? How is it found?
 - What is the difference between a strong acid and a weak acid? Between a strong base and a weak base?
 - What is the difference between an acid and a base?
 - What is a precipitate, and when does it form?
 - What is a limiting reactant? How does it affect how much product is made?
 - What is the vapor pressure of a liquid? Explain how temperature affects this property.
 - Break up the following strong electrolytes.
 - CaCl_2
 - $\text{Sr}(\text{NO}_3)_2$
 - Na_2SO_4
 - K_3PO_4
 - If you needed to make a 0.250 M solution of NaOH that was 500 mL, how many grams of NaOH would you need?
 - If 62.0 g of ZnCl_2 is added to 0.100 L water, what is the molarity of the solution? What is the % (W/V)?
 - If you had 0.025 L of a 0.125 M solution of K_2CrO_4 and heated it until all the water was gone, how many grams of K_2CrO_4 would you have? What was the % (W/V) of the original solution?
 - You have 500 mL of a 12.5% (W/V) solution of NaOH. What is the molarity of the solution?
 - If you need 25 mL of a 1.5 M solution of NaCl, what is the % (W/V)?
 - Find the pH for the following if the concentration of H_3O^+ is equal to:
 - $1.75 \times 10^{-3} \text{ M}$
 - $4.67 \times 10^{-4} \text{ M}$
 - $8.92 \times 10^{-5} \text{ M}$
 - 0.00175 M
 - 0.0000475 M
 - $1.5 \times 10^{-12} \text{ M}$
 - Find the $[\text{H}_3\text{O}^+]$ for the following.
 - pH = 1.65
 - pH = 7.95
 - pH = 5.67
 - pH = 12.75
 - pH = 3.25
 - Finish each equation and indicate which are insoluble (a precipitate) with an (s) and which are soluble with an (aq). If there is not an insoluble product, write "no reaction" next to the equation. (Remember to reference the solubility table at the end of the chapter.)
 - $\text{NH}_4\text{Cl} + \text{KBr} \rightarrow$
 - $\text{Pb}(\text{NO}_3)_2 + \text{NaCl} \rightarrow$
 - $\text{KCl} + \text{Na}_2\text{SO}_4 \rightarrow$
 - $\text{CuCl}_2 + \text{Na}_2\text{CO}_3 \rightarrow$
 - $\text{Pb}(\text{NO}_3)_2 + \text{Na}_2\text{CrO}_4 \rightarrow$
 - $\text{Ba}(\text{NO}_3)_2 + \text{Na}_2\text{SO}_4 \rightarrow$
- Use the following equation for problems 19–26.
- $$\text{___ Li} + \text{___ H}_3\text{PO}_4 \rightarrow \text{___ H}_2 + \text{___ Li}_3\text{PO}_4$$
- How many moles of lithium would react with 4 moles of phosphoric acid?
 - How many moles of phosphoric acid would react to produce 0.4 moles of lithium phosphate?
 - If 2.5 moles of phosphoric acid react, how many grams of lithium are needed?
 - If 17.95 g of lithium react, how many grams of lithium phosphate will form?
 - If you want to make 22.2 g of lithium phosphate, how many grams of phosphoric acid would you need to react?

24. If you react 12.5 g of lithium, how many liters of hydrogen will you make at STP? How many liters would you make if it were at 30°C and 1.9 atm?
25. If you react 15.75 g of phosphoric acid and 22.2 g of lithium, how many grams of lithium phosphate will you make?
26. What is the limiting reactant in the last problem?
27. _____ NaOH + _____ HCl → _____ NaCl + _____ H₂O. How many mL of 0.25 M HCl will be needed to neutralize 8.97 mL of 0.15 M NaOH?
28. _____ Sr(OH)₂ + _____ HCl → _____ SrCl₂ + _____ H₂O. What is the concentration of HCl if it takes 32.6 mL to neutralize 25.0 mL of 0.115 M Sr(OH)₂?
29. _____ NaOH + _____ H₂CO₃ → _____ Na₂CO₃ + _____ H₂O. What is the concentration of H₂CO₃ if it takes 14.5 mL to neutralize 57.0 mL of 0.548 M NaOH?
30. _____ NaOH + _____ H₃PO₄ → _____ Na₃PO₄ + _____ H₂O. What is the volume of NaOH needed to neutralize 12.5 mL of 0.275 M H₃PO₄? The NaOH is 0.97 M.
31. Solutions always have a _____ vapor pressure than the pure liquid.
32. Solutions always have a _____ freezing point than the pure liquid.
33. Solutions always have a _____ boiling point than the pure liquid.
34. How would NaNO₃ affect the freezing point of water as compared to Na₂CO₃ if both chemicals are added in equal concentrations? Explain.
35. If the boiling point elevation constant for water at sea level is 0.52°C/mole particles, how many moles of particles are needed to raise the boiling point to 101.6°C? What concentration of CaCl₂ would be needed to raise the boiling point this much?
36. If the freezing point depression constant for water at sea level is 1.85°C/mole particles, what is the freezing point of a 2.0 M solution of NaCl in water?

Solubility Rules

These anions	Form soluble compounds with these cations	Form insoluble compounds with these cations
NO ₃ ⁻ (nitrate)	Most cations	No common cations
CH ₃ COO ⁻ (acetate)	Most cations	Ag ⁺
Cl ⁻ (chloride)	Most cations	Ag ⁺ , Pb ²⁺ , Hg ₂ ²⁺ , Tl ⁺
Br ⁻ (bromide)	Most cations	Ag ⁺ , Pb ²⁺ , Hg ₂ ²⁺ , Tl ⁺
I ⁻ (iodide)	Most cations	Ag ⁺ , Pb ²⁺ , Hg ₂ ²⁺ , Tl ⁺
SO ₄ ²⁻ (sulfate)	Most cations	Ba ²⁺ , Sr ²⁺ , Pb ²⁺ , Ag ⁺ , Ca ²⁺
CrO ₄ ²⁻ (chromate)	Most cations	Ba ²⁺ , Sr ²⁺ , Pb ²⁺ , Ag ⁺
S ²⁻ (sulfide)	NH ₄ ⁺ , cations of column 1, cations of column 2	Most other cations
OH ⁻ (hydroxide)	NH ₄ ⁺ , cations of column 1, and Ba ²⁺ and Sr ²⁺	Most other cations
CO ₃ ²⁻ (carbonate)	NH ₄ ⁺ , cations of column 1 except Li ⁺	Most other cations
PO ₄ ³⁻ (phosphate)	NH ₄ ⁺ , cations of column 1 except Li ⁺	Most other cations